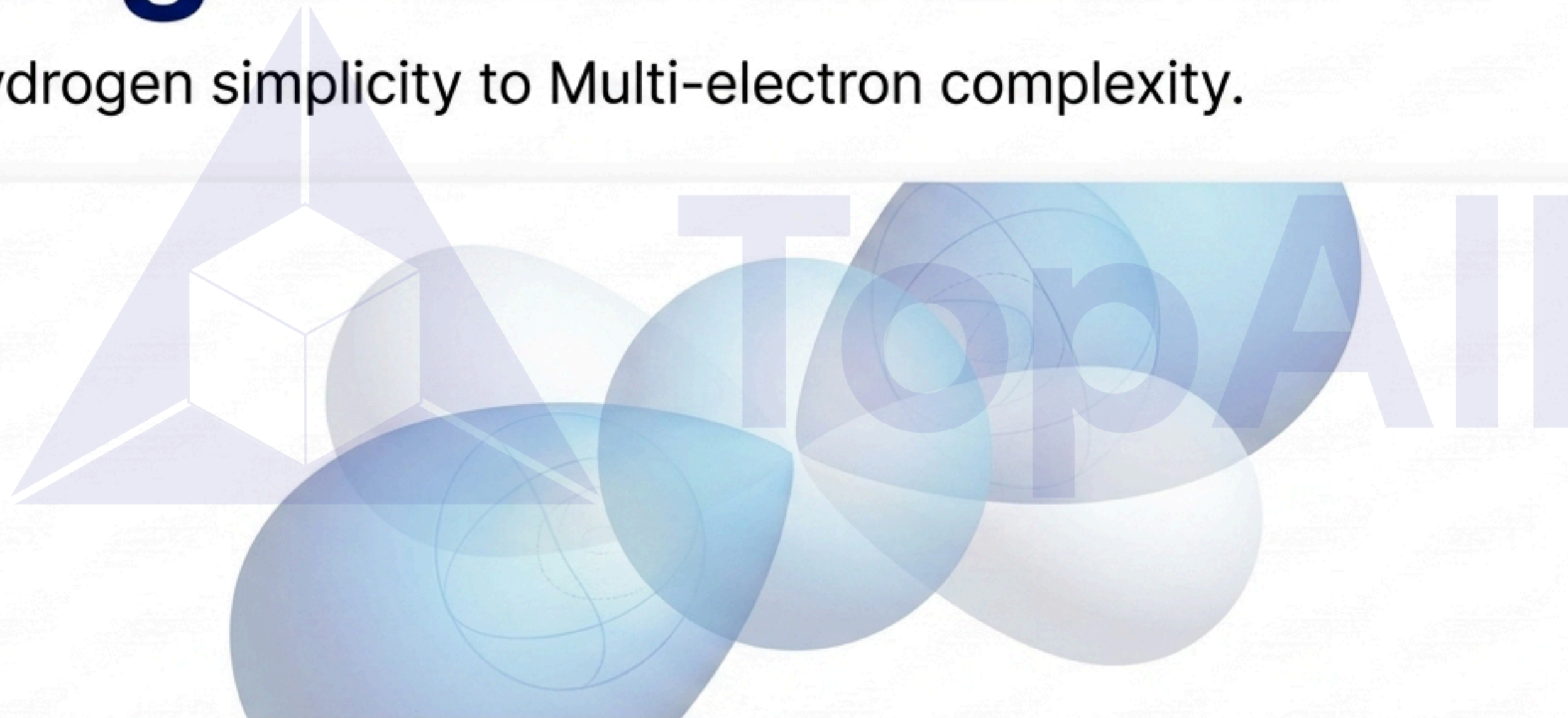


2.6.3 STRUCTURE OF ATOM

Energies of Orbitals

From Hydrogen simplicity to Multi-electron complexity.



[CHEMISTRY CLASS 11]

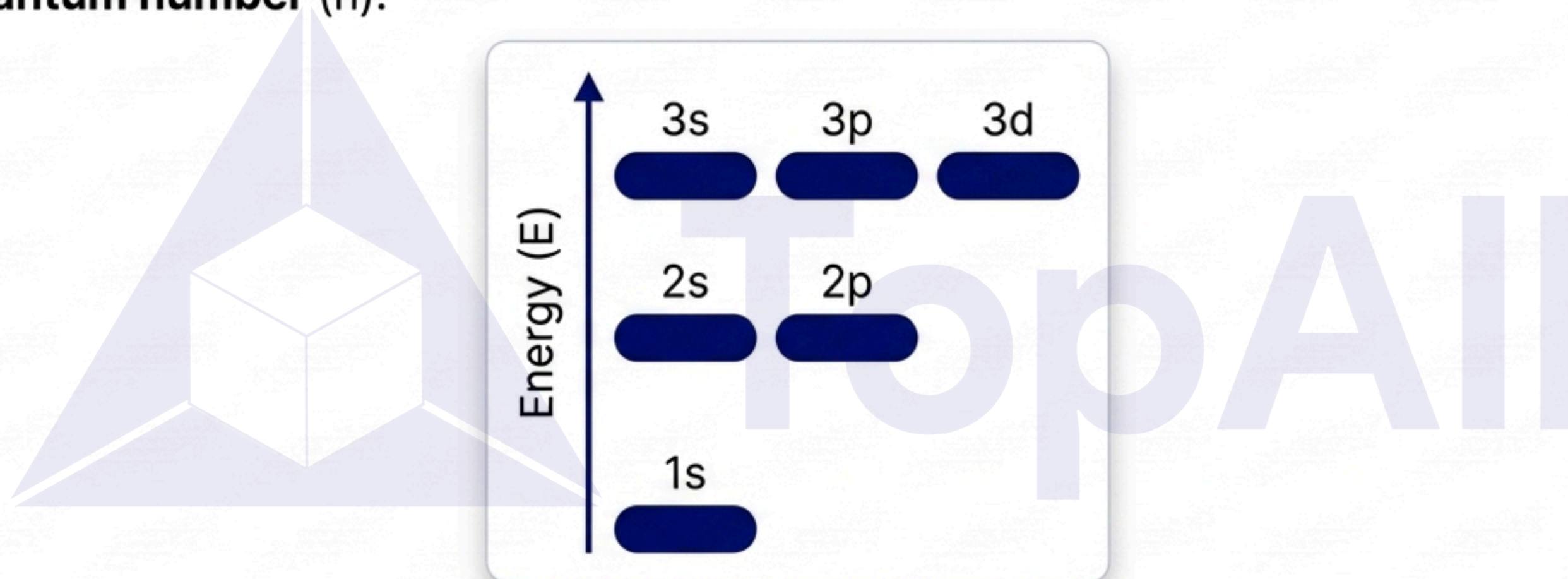
[EXAM PREP]

[QUANTUM MECHANICS]

THE SIMPLE CASE

The Hydrogen Atom Baseline

In a single-electron system like Hydrogen, **the energy of an electron is determined solely by its principal quantum number (n).**



$$1s < 2s = 2p < 3s = 3p = 3d < 4s = 4p = 4d = 4f$$

 As n increases, the orbital size increases, the electron moves further from the nucleus, and the energy becomes less negative (higher).

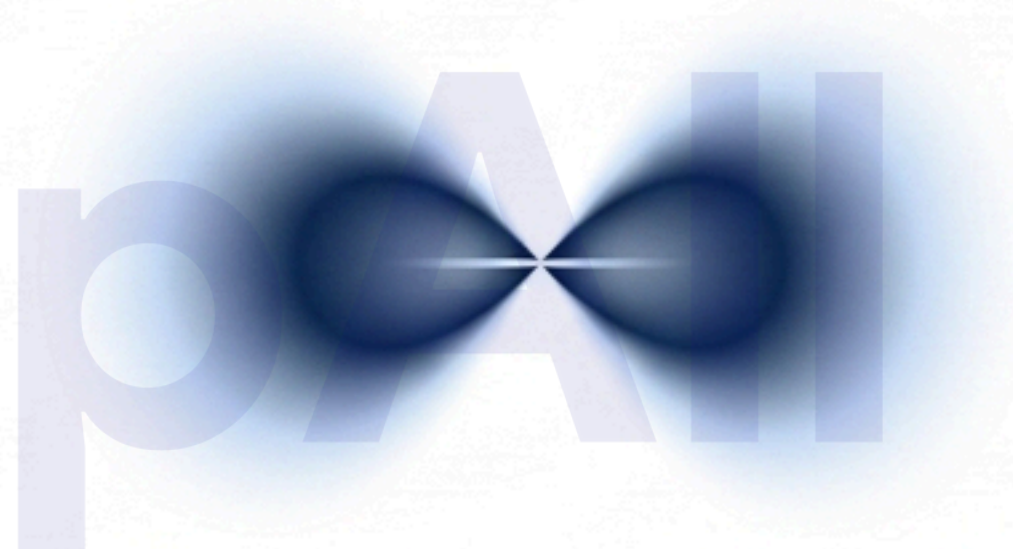
Understanding Degeneracy

Although the 2s (spherical) and 2p (dumbbell) orbitals have fundamentally different shapes, an electron in a Hydrogen atom has the exact same energy regardless of which one it occupies.

[2s Orbital]



[2p Orbital]



=

Definition: Orbitals having the exact same energy are called degenerate orbitals.

In Hydrogen, the 1s orbital is the ground state (most stable). An electron in 2s, 2p, or higher is in an excited state.



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$$E=mc^2$$



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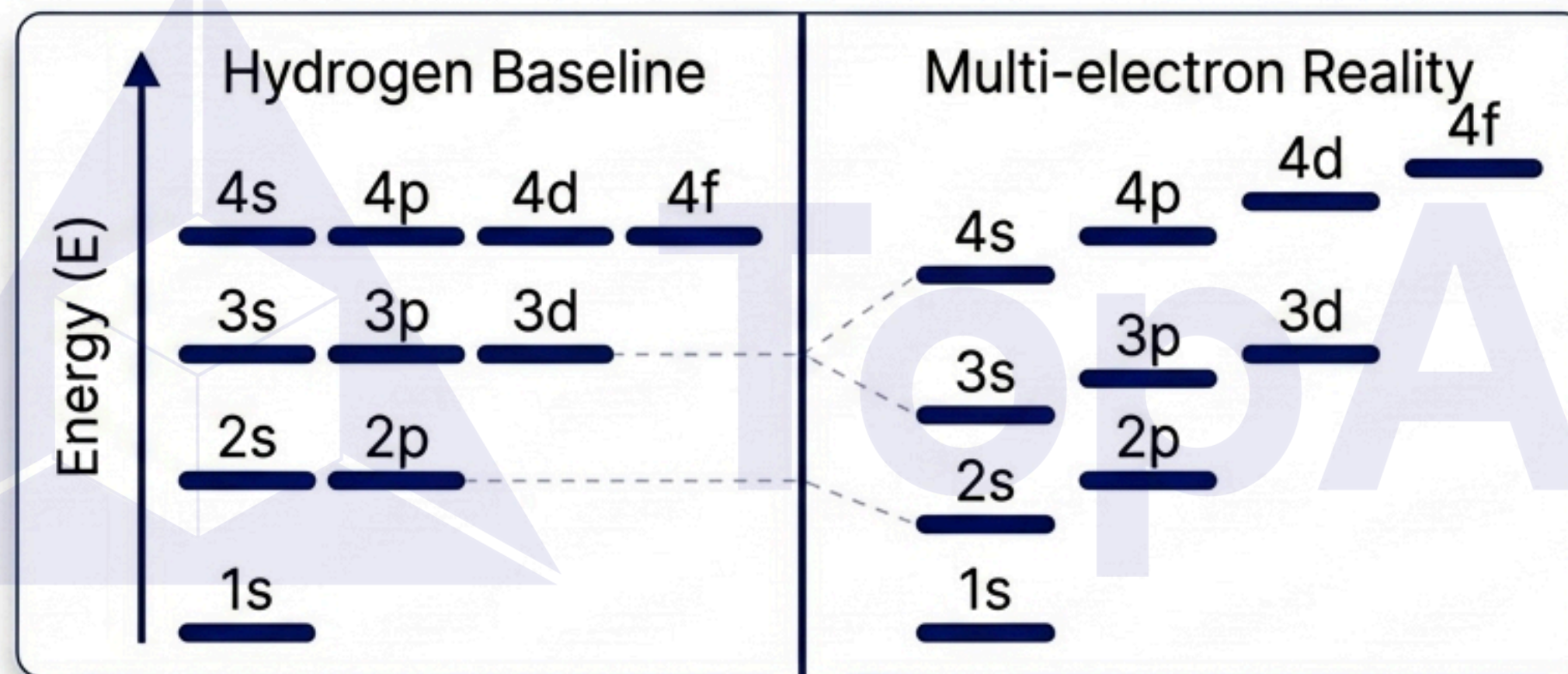
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The Multi-electron Shift

When an large atom has more than one electron, energy depends on both the **principal quantum number** (n) and the **azimuthal quantum number** (l).



Within a shell (same n): $s < p < d < f$

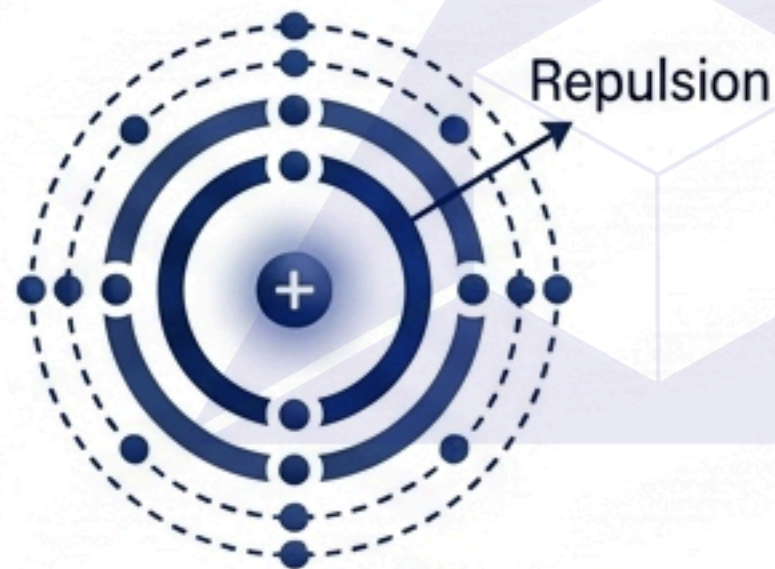
Why the split? Mutual repulsion among electrons creates a staggered energy staircase. Subshells no longer share the same energy.

THE MECHANISM

Why Subshell Energies Split

In multi-electron atoms, inner electrons block outer electrons from feeling the full pull of the nucleus.

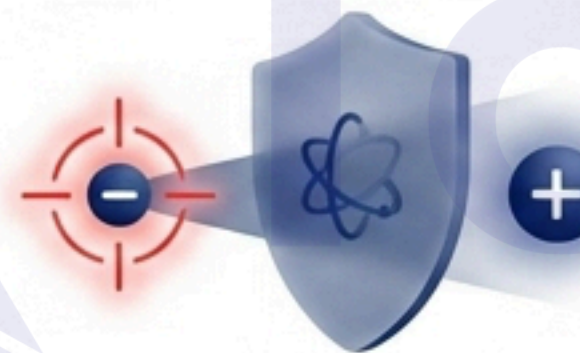
Shielding Effect



Concentric Shield UI

Inner shell electrons act as a screen, repelling outer electrons.

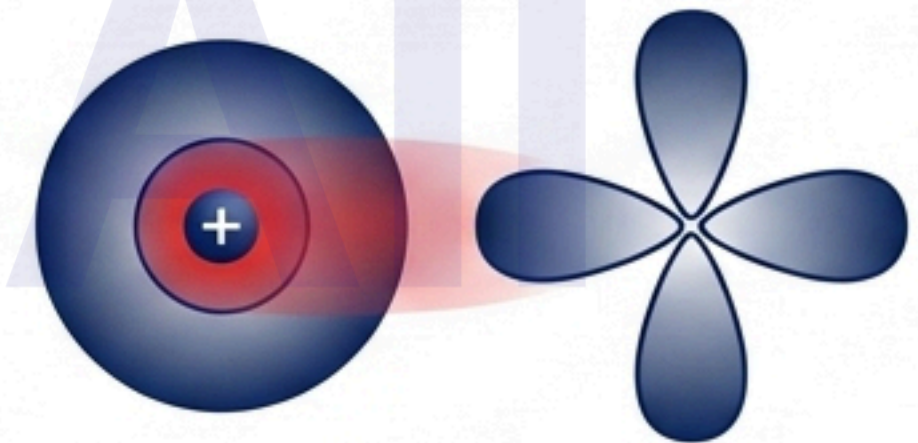
Effective Nuclear Charge (Z_{eff})



$$Z_{\text{eff}} = Z - \text{Shielding}$$

The net positive charge actually experienced by an outer electron.

Penetration Power



Because of its spherical shape, an s-orbital electron spends more time near the nucleus than p, d, or f electrons of the same shell.

Because s penetrates better, it feels more Z_{eff} , making it more tightly bound (lower energy). Therefore: $E(s) < E(p) < E(d) < E(f)$

Predicting Orbital Order

How do we know which orbital fills first in larger atoms? We use the $(n+l)$ rule to navigate the staggered energy levels.

Rule 1

The lower the value of $(n + l)$, the lower the energy.

Rule 2

If two orbitals have the same $(n + l)$ value, the orbital with the lower n has the lower energy.

$$4s: n=4, l=0 \rightarrow (4+0) = 4$$

$$3d: n=3, l=2 \rightarrow (3+2) = 5$$

Conclusion: 4s fills before 3d ($4 < 5$)

Order of filling: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s...

(This energy sequence is the foundation of the Aufbau Principle for electronic configuration).



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